Discussion

This study of the oxidation of alcohols confirms several significant features of the photocatalytic oxidation of organic compounds on ZnO in light which is absorbed only by zinc oxide.

If the initial reaction is the transfer of an electron from zinc to O_2 adsorbed on the surface, the resulting $\cdot O_2^-$ is a radical ion, and evidently requires water or a highly polar solvent for the efficiency of the electron transfer process. The reaction does occur to some small extent in absolute methyl alcohol, but not at all in pure glycerol; *cf.* Fig. 1.

Organic additives undergo oxidation either by electron transfer to the photopositive zinc oxide surface (after loss of a photo-excited electron by the crystal to O_2) or by hydrogen abstraction by the peroxide radical ion to form $HO_2\cdot^-$. In the case of the phenols the phenolate ion adsorbed on the zinc oxide surface can transfer an electron directly to the crystal and then undergo immediate destructive oxidation by the peroxide radical ion or by molecular oxygen. This interpretation would explain the unique ability of the phenols to promote rapid accumulation of hydrogen peroxide at very low phenolic concentrations; *cf.* ref. 1.

In the case of the aliphatic alcohols the rate of oxidation is evidently related to the case of abstraction of hydrogen atoms: primary alcohols, in general, are more easily oxidized than the isomeric secondary or tertiary alcohols. Since these alcohols are probably not initially adsorbed on the ZnO in an ionic form, it is not surprising that higher concentrations are required for rates of peroxide formation equivalent to that of the phenols. The fact that tertiary alcohols are oxidized appreciably is a further indication of the non-specific character of the initial oxidation step (*cf.* ref. 2.)

Primary alcohols would be expected to form radicals of the type $R\dot{C}(H)(OH)$. These radicals can easily lose a second hydrogen atom to form an aldehyde, or react with molecular oxygen to form a peroxide radical, $RC(H)(OH)O_2\cdot$. Subsequent reaction of this radical with a molecule of the original alcohol may lead to hydroxyalkyl hydroperoxides and their derivatives, thus initiating chain reactions. Hydroxyalkyl hydroperoxides dissociate in water in dilute solutions to form aldehydes and hydrogen peroxide; the dissociation is catalyzed by dilute acids and bases, and so would be highly probable under the conditions used for testing for peroxides.[5] Differences in rates of propagation of peroxide chains and differences in rates of hydrolysis of intermediate oxidation products are probably responsible for the variation in rates of peroxide formation in the isomeric alcohols tested. Figures 2 and 3 show that formaldehyde itself is rapidly oxidized in this system; therefore it is not likely that aldehydes will be the final products.

In the case of the higher aliphatic alcohols there are evidently fairly stable organic peroxides present in the irradiated suspensions. They react only slowly with the iodide ion; oxidation products of hexyl and heptyl alcohols continue to release iodine over the course of one or two hours. Tests for peroxide with $Ti(SO_4)_2$ give only about half the peroxide values obtained with iodide. The rest of the peroxide, not reacted with $Ti(SO_4)_2$, must be tied up in fairly stable addition products of the type RCH(OH)(OOH), or possibly as bis-hydroxyalkyl peroxides. Recent work of Mosher, *et al.*, indicates an increasing stability toward thermal decomposition on the part of the higher hydroxyalkyl peroxides. It is possible that they are more stable toward hydrolysis and photochemical decomposition also.

There must be some explanation of the increase in initial rate of oxidation of these higher aliphatic alcohols. It seems likely that their effect on the surface tension of water makes for better contact between oxygen, water and alcohol on the surface of the ZnO. The foaming of 0.01 *M* solutions of hexyl, heptyl and benzyl alcohols on shaking with air and water is quite noticeable, and it is also noticeable that zinc oxide seems to be held in more homogeneous suspension with the higher aliphatic alcohols than it is in the case of the phenols and the lower alcohols.

Acknowledgment.—This work was made possible in part by a grant from Sigma-Xi-Resa Research Fund.

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[5] A. V. Tobolsky and R. B. Mesrobian, "Organic Peroxides," Interscience Publishers, Inc., New York, N. Y., 1954, pp. 44, 45.

[6] C. F. Wurster, L. J. Durham and H. S. Mosher, THIS JOURNAL, **80**, 327 (1958).

WEST HARTFORD 7, CONN.

[CONTRIBUTION FROM THE GEORGE HERBERT JONES LABORATORY, DEPARTMENT OF CHEMISTRY, UNIVERSITY OF CHICAGO]

The Entropies of Activation in the Nitration of Ethyl Benzoate[1]

BY WM. J. LE NOBLE[2] AND G. W. WHELAND

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The nitration of ethyl benzoate by nitric acid in acetic acid as the solvent, and with sulfuric acid as a catalyst, was studied over a range of 100°. The isomer distribution was determined by means of the isotope dilution method. The results show that the three simultaneous substitution reactions have nearly the same entropy of activation, so that the differences between the three rates of formation can be almost entirely ascribed to the differences in activation energy.

The isomer distributions in the disubstitutions of benzene have been the subject of many investigations and a semi-quantitative understanding of the problem has been reached. In the theoretical treatments, the most significant factor affecting this distribution is commonly presumed to be the differences between the potential energies of the

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[1] Based on the dissertation submitted by W. le Noble in partial fulfillment of the requirements for the Ph.D. degree.

[2] N.S.F. Fellow, 1955–1957.

several transition states, and hence the differences between the activation energies of the several corresponding reactions. Variations in the entropy of activation are accordingly assumed to be negligible.[3] Steric effects have been observed to influence the distributions of isomers by decreasing the yields of the *o*-disubstituted products when either the substituents that are initially present or the attacking reagents are bulky[4]; however, since such steric effects manifest themselves in the energy or in the entropy of activation, or in both, they need not here be separately considered.

Although, as was noted above, the entropy of activation in an aromatic substitution reaction is generally considered to be independent of the point of attack, there is apparently no conclusive experimental support for this view. In 1913, to be sure, Scheffer[5] published a paper bearing on the subject; he quoted data taken from the work of Holleman and co-workers, and he showed that the ways in which the observed ratios of isomeric products varied with the temperature in several substitutions were consistent with the belief that, for a given pair of reactants, the entropy of activation is independent of the point of attack. The reliability of his conclusion may be questioned, however, on the following grounds. The analyses by Holleman, *et al.*, were crude by present-day standards and often applied to only two of the three isomeric products; several of the reactions were studied at only two temperatures, and the temperature was not always the only variable; apparently, many of the reactions were not carried out in homogeneous media. For these reasons, it appeared desirable to study some particular reaction with greater care. When the present work was well under way, Brown and McGary[6] published the results of their study of the mercuration of toluene; the entropies of activation for the three isomeric products of that reaction were found to be equal within 0.5 e.u.

Discussion and Results

The nitration of ethyl benzoate was chosen for investigation, since the products formed from this compound would probably not be subject to further attack (see below), and since the reaction had been previously studied by Ingold and Smith.[7] These authors found that, if the reaction is carried out in acetic anhydride at 18°, the product consists of 24% of the *ortho*, 72% of the *meta* and 4% of the *para* isomers. Acetic anhydride could not be used as the solvent in this work, however, since its oxidation by nitric acid, rather than the nitration of the ethyl benzoate, becomes the main reaction at higher temperatures. Acetic acid therefore was used. With nitric acid, this substance forms a compound[8] ($2CH_3CO_2H \cdot HNO_3$) which boils at 127°, and which is suitable as a nitrating agent since, within the temperature range of interest, it remains liquid and does not decompose; these latter favorable properties are retained also when ethyl benzoate and sulfuric acid are added. Nitrous acid (or some material which is estimated as nitrous acid) has been shown to retard the nitration of deactivated benzene derivatives.[9] This effect seems to be due to the removal of nitronium ion by nitrous acid, so that the isomer distribution is presumably not affected; it is simpler, however, to eliminate the impurity than to prove that purification is unnecessary. The order of mixing the acetic acid, the nitric acid and the catalyst (sulfuric acid) is important if the concentration of nitrous acid is to be kept low. The procedures for preparing the solvent and for estimating the nitrous acid present are described below (see Experimental). Sulfuric acid was employed as a catalyst because, without it, the reaction is inconveniently slow. In a study of the approximate rate of disappearance of ethyl benzoate, ultraviolet spectra were used for analysis. The success of this procedure was dependent on the fact that, at 2600 Å., the optical density of unsubstituted ethyl benzoate is much less than that of any of its mononitrated derivatives; with a path length of 1 cm., the optical densities of $10^{-4}M$ solutions of ethyl benzoate, ethyl *o*-nitrobenzoate, ethyl *m*-nitrobenzoate and ethyl *p*-nitrobenzoate are, respectively, 0.095, 0.384, 0.707 and 1.31. The distribution of isomers in the product was taken from the more precise analyses described below. With the use of these data, the rate of nitration at 85° was found to be approximately 50 times as great in the solvent described in the Experimental part as in a solvent differing only in the omission of the sulfuric acid catalyst. These tentative studies showed that the effective value of the entropy of activation for the reaction as a whole was roughly 40 e.u.

Two possible side reactions had to be tracked down. First of all, there was the possibility of dinitration. Ultraviolet analysis showed, however, that ethyl *m*-nitrobenzoate, for example, produced no detectable amount of ethyl 3,5-dinitrobenzoate under the most extreme conditions used in the determination of the isomer distribution. A somewhat more difficult problem was that of transesterification. Ordinarily, this term is used to indicate the reaction of an ester with an alcohol. However, a similar reaction may take place between an ester and an acid.[10] If the ethyl benzoate is partly transesterified to yield benzoic acid and ethyl acetate, the resulting benzoic acid will react to form the nitrobenzoic acids, but, if these acids are separated from the esters before analysis, this type of interference is not serious. On the other hand, if the ethyl nitrobenzoates themselves are also partly converted into the corresponding acids, these fractions would also be lost in the isolation of the esters, so that the subsequent assay would be in error. Finally, it is conceivable that, if the various species that are present are in mobile equilibrium with one another, the mixture of isomeric nitrated esters isolated for

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[3] L. P. Hammett, "Physical Organic Chemistry," McGraw-Hill Book Co., Inc., New York, N. Y., 1940, p. 124.

[4] H. C. Brown and W. H. Bonner, THIS JOURNAL, **76**, 605 (1954).

[5] F. E. C. Scheffer, *Proc. Acad. Wet. Amsterdam*, **15**, 1118 (1912-1913).

[6] H. C. Brown and C. W. McGary, Jr., THIS JOURNAL, **77**, 2306 (1955).

[7] C. K. Ingold and M. S. Smith, *J. Chem. Soc.*, 905 (1938).

[8] A. Pictet and P. Genequand, *Ber.*, **35**, 2526 (1902).

[9] E. D. Hughes, C. K. Ingold and R. I. Reed, *J. Chem. Soc.*, 2400 (1950).

[10] W. B. Renfrow and A. Chaney, THIS JOURNAL, **68**, 150 (1946).

analysis may have been derived in part from benzoic acid, and not exclusively from ethyl benzoate. Since sulfuric acid, which was introduced in order to accelerate the nitration, also catalyzes this type of transesterification, an investigation seemed desirable. For the estimation of the extent to which transesterification competes with the nitration, each of the three product esters was exposed to the nitrating mixture. At regular intervals, samples of the solution were withdrawn and distributed between water and 2,2,4-trimethylpentane, and each phase was then analyzed for its aromatic content by means of its ultraviolet absorption. The distribution coefficients both of the esters and of the acids had previously been measured by the same spectroscopic methods, and so the rate of transesterification of each could be determined. The experiments at 85°, conducted over a prolonged period of time, showed that the transesterifications of the ethyl nitrobenzoates are slow. This method is not applicable to ethyl benzoate, which is rapidly consumed by nitration, but the product of nitration was found to contain no significant amount of the free nitrobenzoic acids. The conclusions reached on the basis of these studies are that the transesterification of ethyl benzoate does not seriously compete with the nitration, that the ethyl nitrobenzoates arise solely from nitration of ethyl benzoate, and that the losses of these products by transesterification are small. Appropriate corrections were made in the experimental results to account for these losses.

Analysis of the reaction mixture presented a major problem, for the precision of the rate ratios is of course limited by the precision with which each isomer can be estimated. Furthermore, the solutions to be assayed must be dilute; for, if too much ethyl benzoate is used for the reaction, the heat of reaction makes it difficult to keep the temperature constant. Because of the necessarily low concentrations, spectral analysis could not be used. The method used by Ingold and Smith[7] likewise is not very precise. It seemed therefore that the isotope dilution techniques offered the best solution. The method can be applied in either of two ways. In the first, a known amount of a labeled product is added to a known fraction of the reaction mixture; some of this compound is then isolated and examined for its isotope content. The dilution factor cannot be very large here, since otherwise the diluent and the isolated products are too similar in isotope contents. This method was rejected when it was found that a procedure for separating the aromatic acids reported by Griess[11] could not be repeated in this Laboratory. In the alternative method, which was the one adopted, the reactant is the labeled material. For reasons of economy and of convenience of preparation, the best isotopes for this work seemed to be those of hydrogen. A labeled hydrogen atom in the ethyl group has the advantage of a simple preparation (esterification of C_6H_5COOD with diazoethane), but this procedure requires separation of the three esters; a ring-bound hydrogen isotope was therefore preferred. Isotope analysis then can be performed on the acids which are more easily purified.

Before the isomer distributions could be investigated, the following points had to be considered: (1) Preparation of labeled ethyl benzoate: since calculations are much simpler if they can be based on identical isotope contents in the three positions, it was desirable to prepare ethyl benzoate statistically labeled in the ring. The most obvious method, consisting of exchange of benzene with D_2SO_4, followed by bromination, Grignard reaction and esterification affords only a rather low yield with respect to the isotope. A direct method for labeling benzoic acid in statistical manner, which was worked out recently by Garnett[12] in this Laboratory, was therefore used. (2) Study of possible exchange during nitration and during hydrolysis of the esters: by the use of D_2SO_4 as a catalyst for the nitration and of D_2O during hydrolysis, it could be shown that none of the esters contained any deuterium, so that these exchange reactions evidently did not occur. (3) Study of the isotope effect: Melander's work[13] indicated that a significant isotope effect is unlikely but, since our results should be very sensitive to such an effect, it was decided to conduct the experiment with both deuterium and tritium lables in order to screen it out if it occurred. (4) Purification of the nitrobenzoic acids: the results are again very sensitive to highly labeled impurities, so that considerable work was required to show that the analyses were performed on products of sufficient purity. The purification consisted of repeated recrystallizations from several solvents. The criterion of purity of the labeled acids was the depression of their isotope contents upon scavenging with the possible unlabeled impurities. No such depression was observed if the procedure described in Experimental was followed.

During the nitration experiments the temperatures rose by as much as 2° at the highest temperatures. Observation of the temperature as a function of time permitted calculation of an average temperature accurate to about 0.1°. The tritium contents were measured by the technique described by Wilzbach, Kaplan and Brown[14]; deuterium analyses were carried out by a modified version of the method of Graff and Rittenberg[15] (see Experimental). The results are summarized in Table I and in Fig. 1. The smaller precision of the results that involve the rate at the ortho positions is due to difficulties in purifying *o*-nitrobenzoic acid. The differences between the lines obtained for deuterium and tritium are due to inherent small differences in both precision and accuracy of the standard methods for assaying these isotopes. In the calculations allowance has been made for the considerations that there is only one *p*-position, and that 20% of the isotope is removed during the reaction.

In Fig. 1, the lines should be straight and they should be horizontal if the entropy of activation is the same for all points of attack, since *e. g.* (ref. 3, p. 120)

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(11) P. Griess, *Ber.*, **8**, 526 (1875).

(12) J. L. Garnett, Ph.D. Thesis, The University of Chicago.

(13) L. Melander, *Nature*, **163**, 599 (1949).

(14) K. E. Wilzbach, L. Kaplan and W. G. Brown, *Science*, **118**, 522 (1953).

(15) J. Graff and D. Rittenberg, *Anal. Chem.*, **24**, 878 (1952).

$$\Delta S^{\ddagger}_{m} - \Delta S^{\ddagger}_{p} = \frac{\partial RT \ln (k_m/k_p)}{\partial T}$$

When straight lines are fitted by the method of least squares to the experimental points of Fig. 1, the derived entropy terms for substitution at the *o*-, *m*- and *p*-positions, respectively, are not quite identical. When deuterium was used as the tracer, $\Delta S_0^{\ddagger} - \Delta S_m^{\ddagger}$ was found to be 1.8 ± 0.4, $\Delta S_0^{\ddagger} - \Delta S_p^{\ddagger}$ was found to be 0.4 ± 0.4 and $\Delta S_m^{\ddagger} - \Delta S_p^{\ddagger}$ was found to be −1.4 ± 0.2 cal./degree/mole; when tritium was used as the tracer, the corresponding values were found to be 0.8 ± 0.6, 1.4 ± 0.6 and 0.6 ± 0.2, respectively (the anomalous values obtained at −4.0° were shown to result from an error in the isotope analysis; they were not included in the least squares treatment), and finally, in the approximate rate studies that were described above, the effective value of the entropy of activation for the nitration as a whole was found to be roughly 40 e.u. The margins indicated are standard deviations. The results would suggest that the quantities $\Delta\Delta S^{\ddagger}$ differ in nearly every case from zero by a small but definite amount, and that $\Delta S_m^{\ddagger} - \Delta S_p^{\ddagger}$ is definitely dependent on the isotope used. However, the latter difference (of about 0.5 e.u.) may well have been caused by nonrandom errors, such as a small deviation from true statistical labeling in the two starting materials, or a small difference in accuracy for the two methods of isotope assay. It does not seem justified therefore to ascribe this dependence to any unusual isotope effect unless more accurate information becomes available. The results are the best evidence which is now available to support the view that there are no large differences between the entropies of activation for nitration at the various positions in an aromatic compound.

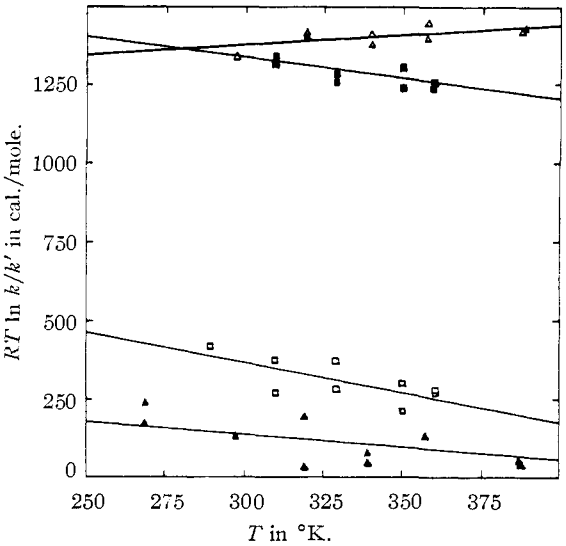

Fig. 1.—Rate ratios in the nitration of ethyl benzoate: △, *meta/para* with ethyl tritiobenzoate; ■, *meta/para* with ethyl deuteriobenzoate; ▲, *meta/ortho* with ethyl tritiobenzoate; □, *meta/ortho* with ethyl deuteriobenzoate.

Experimental

Tritium Analysis.—The method of Wilzbach, Kaplan and Brown[14] was used without change.

TABLE I

RATE RATIOS IN THE NITRATION OF ETHYL BENZOATE

Isotope used	Temperature, °C.	ln (*meta/para*)	ln (*meta/ortho*)
D	16.25	. .	0.73
D	36.44	2.14	.44
D	36.43	2.18	.61
D	55.87	1.96	.43
D	55.80	1.92	.57
D	77.11	1.78	.30
D	77.03	1.87	.43
D	87.45	1.73	.37
D	87.35	1.75	.39
T	−4.0	1.61	.32
T	−4.0	1.51	.45
T	24.27	2.27	.22
T	46.05	2.21	.31
T	46.09	2.23	.05
T	66.69	2.04	.07
T	66.53	2.09	.12
T	84.49	2.03	.19
T	84.43	1.96	. .
T	114.5	1.84	.07
T	115.5	1.85	.05

Deuterium Analysis.—The method of Graff and Rittenberg[15] was followed with these modifications: the water-jet pump was replaced by a Duo-Seal vacuum pump capable of maintaining a vacuum at one micron, and the McLeod gage was replaced by a continuously reading Weston vacuum gage. Since there is a considerable isotope fractionation effect in the reduction of water by zinc,[16] it was necessary to reduce at least 99.9% of the water to hydrogen to ensure 1% accuracy. This could not be achieved conveniently by the apparatus described by Graff and Rittenberg.

The readings obtained from the mass spectrometer differed considerably from those calculated from standard samples prepared by mixing deuterium oxide and water, particularly for low deuterium concentrations, so that the use of standards was imperative. In calculations, the natural deuterium content (0.016%) of the diluting material was taken into account at all times. In a typical analysis, a sample of 3-5 mg. of compound (which need not be weighed) was entered in a porcelain boat into a quartz combustion tube and burned there in a stream of oxygen which had been thoroughly dried by passing through concentrated sulfuric acid, over magnesium perchlorate and over zinc granules at −80°. The combustion gases were passed through an oven containing copper oxide maintained at 750°, and then through a trap kept at Dry Ice temperature. After the burning was completed, the oxygen supply was shut off and the combustion tube was evacuated until a pressure of 3–5 μ obtained; the water was isolated in the trap. This water was then allowed to evaporate and pass through an oven containing zinc, which was carefully maintained at a temperature of 380–400°. At lower temperatures the reduction was inefficient; at higher temperatures the zinc would melt and clog the oven. Small adjustments of the temperature could be made *via* a circuit involving two Variacs in series. The oven was carefully protected against air drafts, as local cooling caused rapid sublimation of the zinc and clogging of the oven. The hydrogen was forced by a Toepler pump into a collector tube; remaining traces of water were frozen out in a second trap. This water was again passed through the oven, and so on. Since the total hydrogen pressure was usually about 10 mm., the reductions and Toepler operations were continued until the hydrogen pressure in the line had decreased to about 10 μ. Usually the reduction was complete in two passes. Four samples of each compound were burned, the first two being discarded, since the glass walls and the zinc oven retained some hydrogen so that there was a considerable memory effect. Isotope ratios were found on a Consolidated mass spectrometer by reading currents on mass two and mass three peaks. Standards in the same range of deuterium content were used for comparison.

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(16) W. G. Henderson, Jr., and R. B. Bernstein, THIS JOURNAL, **76**, 21 (1954).

Ultraviolet Analysis.—Ultraviolet spectra of the four aromatic acids and their ethyl esters in several solvents were measured on a Beckman recording spectrophotometer. Beer's law was found to be obeyed in all cases of interest in this work.

Preparation of Nitrating Solvent.—Nitric acid was purified by mixing equal volumes of concentrated sulfuric acid and 70% nitric acid, and slowly evacuating to 1 mm. pressure. The mixture thereby cooled to about $-5°$. The nitric acid was frozen in a liquid air trap. For the second distillation the distillate was mixed with about 10% of its volume of concentrated sulfuric acid. This process was repeated until the distillate was colorless; usually four distillations sufficed. One mole of this nitric acid slowly was poured on 2.01 moles of frozen acetic acid; the liquefied mixture was refrigerated at 0° for five days; 96% sulfuric acid (10% of the volume of the original mixture) was then added and the entire mixture was heated in an oil-bath at 83–85° for 30 minutes.

Estimation of Nitrous Acid.—The procedure used by Ingold[17] proved convenient. A 1-ml. sample solution was mixed with 10 ml. of water, 1 ml. of 1% aqueous sulfuric acid and 1 ml. of a saturated aqueous solution of sulfanilic acid. This mixture was allowed to stand for 10 minutes. Then 1 ml. of 1% aqueous α-naphthol was added. This mixture was diluted to 25 ml. with concentrated ammonium hydroxide and allowed to stand for 24 hours. Then the resulting red dye was estimated colorimetrically by comparison with a similar solution obtained with an approximately 10^{-5} *M* sodium nitrite solution. Various procedures for mixing acetic acid, nitric acid and sulfuric acid were tested in this manner. Small traces of nitrous acid from these mixtures could be shown by this analysis to be removed by heating at 85° for 30 minutes.

Approximate Rate of Nitration of Ethyl Benzoate.—Two small flasks, each containing 50 ml. of the nitrating solvent, were kept at 85° in an oil-bath, the temperature being manually maintained by means of a Variac. Ethyl benzoate (0.1 ml.) was introduced into one of the flasks and at regular intervals 200-mg. samples were removed from each flask by means of a micropipet. The samples were transferred to 10-ml. volumetric flasks containing methanol. After being weighed, the flasks were filled to the mark with methanol. Analysis was carried out by ultraviolet spectroscopy, the nitrating solvent obtained from the corresponding blank sample serving as solvent for the reference cell. A similar experiment was conducted at 20°.

Nitration of Ethyl *m*-Nitrobenzoate.—The experiment, conducted at 85°, was entirely similar to that described in the preceding section except that 0.1 g. of ethyl *m*-nitrobenzoate was used instead of ethyl benzoate. Samples were taken for 5 hours and ultraviolet spectra were determined. Although the ultraviolet spectra of ethyl *m*-nitrobenzoate and ethyl 3,5-dinitrobenzoate are quite different, no change in the spectrum of the solution could be detected.

Transesterification Experiments.—Ultraviolet spectra of the acids and the ions in water were recorded. By means of these spectra and those obtained for the esters in 2,2,4-trimethylpentane it was possible to determine distribution coefficients of each of the esters and the acids between water and 2,2,4-trimethylpentane. In experiments at 20 and 85°, similar again to those described above, 0.1 g. of an ester was introduced into 50 ml. of the nitrating solvent. Samples were distributed between 5 ml. of water and 5 ml. of 2,2,4-trimethylpentane and ultraviolet spectra of both phases were measured. Knowledge of the distribution constants then allowed evaluation of the extent to which transesterification had occurred. During one half-life of ethyl benzoate in the nitrating solvent at both 20 and 85°, these amounts of transesterification were found: ethyl *o*-nitrobenzoate, 0.1%; ethyl *m*-nitrobenzoate, 0.3%; ethyl *p*-nitrobenzoate, 0.3%.

Preparation of Labeled Ethyl Benzoate.—For the preparation of the tritiobenzoate, one hundred mC. of tritium gas, obtained from the United States Atomic Energy Commission, was passed over copper oxide at 400° in a vacuum line, the water formed being frozen out in a trap. That part of the line which had been in contact with the tritium was flushed repeatedly with hydrogen, which also was oxidized and collected as water in the trap. The water was distilled into a tube containing several grams of benzoic acid and about 5% of its weight of platinum oxide which had previously been reduced by hydrogen. The water in this tube was frozen and the tube was sealed under vacuum. The tube was subsequently shaken in an oven at 130° for 72 hours. After the tube had been cooled by liquid nitrogen, it was opened and the contents were removed by ether. Tritium in the water and in the carboxyl group of the benzoic acid was removed carefully by extraction with excess water, traces of benzoic acid being removed from the water washings by shaking with ether. The ether was distilled off, and the residue was dried in a vacuum desiccator overnight. The acid was then dissolved in ether and esterified with diazoethane. After appropriate dilution by natural ethyl benzoate the product was purified by distillation.

The preparation of ethyl deuteriobenzoate was entirely similar, except that deuterium was available as pure deuterium oxide. The procedure results in statistical labeling of the aromatic ring.[12]

Determination of Isomer Distribution in the Nitration of Ethyl Benzoate.—Three hundred ml. of the nitrating solvent, prepared as described previously, was introduced into a 500-ml. 3-neck flask fitted with a mechanical stirrer with ground glass seal; the seal was water cooled and lubricated with a few drops of acetic acid. One neck was fitted with a gas-inlet tube and the other with a tube reaching to the bottom, so that nearly all of the contents could be blown out. The flask also had a thermometer well containing some mercury for good heat conductance. A long-stem 0–100° precision thermometer was used; it allowed readings to 0.01°. The flask was nearly entirely immersed in a large bath containing mineral oil. The oil was heated by two knife heaters of 250 w. each. The temperature was adjusted and controlled by the usual circuit of a mercury thermoregulator, battery and relay. Usually one hour was required for the temperature to reach equilibrium. Three ml. of labeled ethyl benzoate, distilled just before use in a molecular still, then was introduced and the temperature was read as a function of time. When the reaction was about half over, the solution was blown by nitrogen into about 500 g. of crushed ice within a 1-liter volumetric flask. The flask was shaken until the ice had melted; then it was filled to the mark. Portions of this solution were then transferred to smaller volumetric flasks; these were filled to the mark at 0°. Each of these was thereafter treated as follows:

The solution was introduced into a 1-liter separatory funnel; the volumetric flask was rinsed several times with acetic acid. A known amount of natural ester was dissolved in 25 ml. of acetic acid and introduced the same way. After brief shaking, water was poured in to bring the total volume to about 300 ml.; the solution became cloudy as the esters began to precipitate. It was twice shaken with 300 ml. of 30–60° petroleum ether; these washings in turn were combined and washed with 300 ml. of water. The petroleum ether was distilled off; the residue was collected with several washings of 5 ml. of acetic acid. The resulting solution was allowed to evaporate in the hood until the odor of acetic acid was no longer observable. The residue was dissolved again in 3 ml. of acetic acid. The solution was introduced into a thick-walled tube of about 20-ml. capacity. Subsequently 7 ml. of water and 1 ml. of concentrated sulfuric acid were added. This mixture was frozen in liquid nitrogen; then the tube was evacuated, sealed, heated and shaken for 48 hours in the oven at 130°; this treatment effected complete hydrolysis. The tube was cooled and opened, and the contents were removed by ether. The ethereal solution was washed a few times with water. After evaporation and drying, the purification differed somewhat for the three acids:

o-Nitrobenzoic acid was recrystallized twice from benzene, twice from *n*-butyl ether and twice from water; *m*-nitrobenzoic acid was recrystallized once from benzene and three times from water; *p*-nitrobenzoic acid was recrystallized twice from benzene and three times from water. The final products were stored in a vacuum desiccator until ready for isotope analysis. That these purification processes were efficient was shown in one instance by scavenging each labeled pure acid by small amounts of each of the other acids and by benzoic acid. After repurification there was no depression of the isotope contents.

Exchange Studies.—Exchange of the ring-bound isotope was conceivable in two operations: nitration and hydrolysis. The first possibility was examined as follows: some nitrating solvent was prepared by the use of concentrated deuteriosulfuric acid as a catalyst; this had been prepared by sub-

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[17] G. A. Benford and C. K. Ingold, *J. Chem. Soc.*, 929 (1938).

liming the appropriate amount of sulfur trioxide into pure D_2O. In separate experiments ethyl *o*-nitrobenzoate, ethyl *m*-nitrobenzoate and ethyl *p*-nitrobenzoate were exposed to this solvent at 85° for several hours. After recovery of the esters and hydrolysis, no deuterium could be detected in the acid; this fact showed that no exchange occurred. Ethyl benzoate was exposed to this solvent under similar conditions for 5 minutes. A small amount of ethyl benzoate could be recovered and the corresponding acid also had only a natural deuterium content. The second possibility was checked by hydrolyzing each of the esters under the conditions described before, except that pure D_2O was used. After isolation and purification of the acids they could be shown to have only natural deuterium contents.

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CHICAGO 37, ILL.

[CONTRIBUTION FROM THE STAMFORD LABORATORIES, RESEARCH DIVISION, AMERICAN CYANAMID CO.]

Polarographic Behavior of Unsymmetrical Polyhalogenated Compounds. Polychlorinated Propionitrile and Derivatives[1]

BY W. H. JURA AND R. J. GAUL

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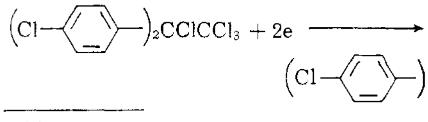
The polarographic behavior of 2,2,3-trichloro- and 2,3-dichloropropionitriles and their derivatives was studied in order to determine the products formed in the reduction of unsymmetrical structures at the dropping mercury electrode. Reduction in all cases results in the formation of the double bond. Polarographic data, substantiated by cryoscopic and nuclear magnetic resonance measurements, are presented which represent the first published experimental evidence for the determination of the *cis* and *trans* isomers of 2,3-dichloroacrylonitrile. Each isomer is reduced to acrylonitrile *via* a different path. Polarograms of 2,2,3,3-tetrachloropropionitrile indicate that it is reduced to *trans*-2,3-dichloroacrylonitrile.

A search of the literature reveals that, while considerable work has been done on the polarographic behavior of halogenated organic compounds, no work has been reported on the behavior of compounds of the general types: $CHX{=}CXY$, CH_2XCHXY, CH_2XCX_2Y, CHX_2CX_2Y, where X = chlorine atom and Y = carboxyl, ester, amide or nitrile.

The closest approach to any of the above structures is found in the work of Elving, *et al.*,[2,3] who made a coulometric and polarographic study of the diastereoisomers of 2,2′-dibromosuccinic acid and their ethyl esters. Briefly, these studies showed that the reduction in all cases involved the elimination of both bromine atoms by a two-electron process to form an olefin. However, both bromine atoms in any of these compounds, excepting the acid at *p*H's where it contains one ionized and one un-ionized carboxyl group, are equivalent since both are alpha to a carboxyl or ester group. Stackelburg[4] also indicated that adjacent halogens could be removed electrolytically to form a stable double bond. Here again, only symmetrical structures were studied wherein the halogens must be considered equivalent.

During the preparation of this manuscript Rosenthal, *et al.*,[5] reported work on 1,1,1,2-tetrachloro-2,2-bis-(*p*-chlorophenyl)-ethane which led them to believe that the first cathodic wave is due to the reaction



Experimental

Apparatus.—All polarograms were obtained with a Leeds and Northrup Electrochemograph, type E, using damping position 1. The $m^{2/3}$, h, and other pertinent values for a given capillary are indicated in the tables. Heyrovsky cells were used throughout this work and polarograms were recorded against an external saturated calomel electrode (S.C.E.) *via* a KCl (aqueous)–KCl/agar/glass plug bridge. In those experiments where alkali metal ions would interfere, the "cell end" of the bridge was washed, dried and inserted into the unstirred solution in the cell just before the polarogram was recorded. A Beckman model G *p*H meter was used for all *p*H measurements. A Serfass conductivity bridge, model RC-M 15, was used to measure the cell resistance of all solutions examined which approached a value of 3000–3300 Ω in solutions containing 30% EtOH–0.05 *M* $(CH_3)_4NI$.

Controlled potential electrolyses were carried out with an automatic apparatus similar to that described by Lingane[6] and the total current passed was measured with a copper coulometer.

Polarographic Procedure.—In most cases, stock solutions of the compound being investigated were made up just before use in anhydrous ethyl alcohol in which all the compounds were stable. Aliquots of these solutions were diluted with predetermined amounts of electrolyte, ethyl alcohol and water. Airco prepurified nitrogen, equilibrated with a portion of the same solution as contained in the cell, was used to remove the oxygen prior to electrolysis. In cases where it was necessary to exceed polarizing voltages of ≃1.9 volts *vs.* S.C.E., deaeration was carried out prior to insertion of the KCl bridge from the S.C.E.

In the case of more volatile compounds, only approximate concentrations were used and the following procedure was adopted. After deaerating 10 ml. of solvent–electrolyte solution in the cell, 0.1 ml. (micro pipet) of a concentrated (∼0.05 *M*) alcoholic solution of the compound in question was added. Nitrogen was passed through the cell for an additional 30 seconds to ensure mixing. After blanketing the solution with nitrogen, the "cell end" of the KCl bridge was inserted and the polarization was carried out.

All $E_{1/2}$ values have been corrected for iR drop and the distances stepped off along the current axis in the figures are equivalent to 1.00 μA. unless otherwise indicated. All potentials cited are "*vs.* S.C.E."

Syntheses.[7] **2,2,3-Trichloropropionitrile.**—2,2,3-Trichloropropionitrile was prepared by the procedure of Brintzinger, *et al.*,[8] with two modifications: the amount of pyri-

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(1) Presented at the Seventh Pittsburgh Conference on Analytical Chemistry and Applied Spectroscopy, Pittsburgh, Pa., February 28, 1956.

(2) P. J. Elving, I. Rosenthal and A. J. Martin, THIS JOURNAL, **77**, 5218 (1955).

(3) P. J. Elving and I. Rosenthal, *ibid.*, **73**, 1880 (1951).

(4) M. V. Stackelburg and W. Stracke, *Z. Elektrochem.*, **53**, 118 (1949).

(5) I. Rosenthal, G. Frisone and R. J. LaCoste, *Anal. Chem.*, **29**, 1639 (1957); private communication.

(6) J. J. Lingane and S. L. Jones, *ibid.*, **22**, 1169 (1950).

(7) All melting points are corrected; boiling points are uncorrected.

(8) H. Brintzinger, K. Phannstiel and H. Koddebusch, *Angew. Chem.*, **A60**, 311 (1948).